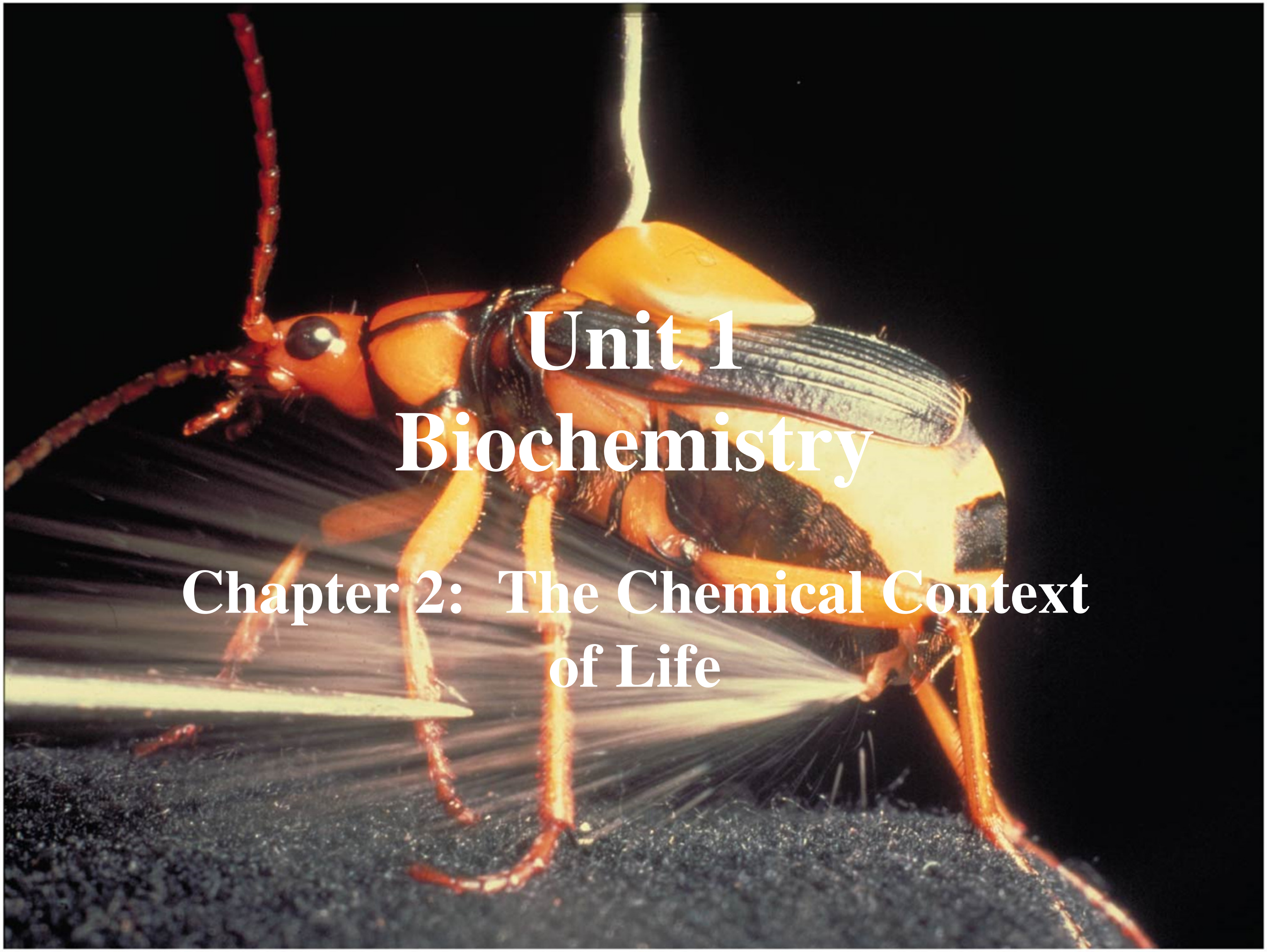


Unit 1
Biochemistry

Chapter 2: The Chemical Context of Life

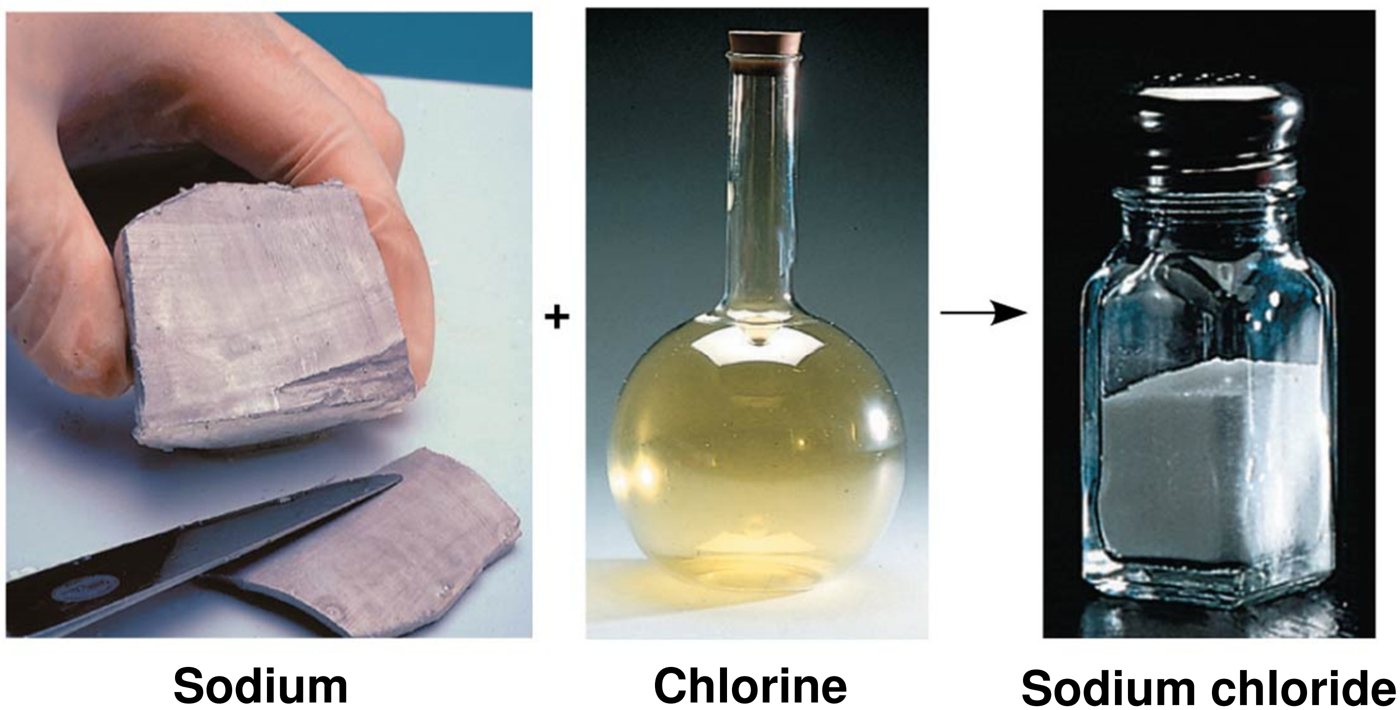
Concept 2.1: Matter consists of chemical elements in pure form and in combinations called compounds

- Organisms are composed of **matter**
- Matter is anything that takes up space and has mass

Elements and Compounds

- Matter is made up of elements
- An **element** is a substance that cannot be broken down to other substances by chemical reactions
- A **compound** is a substance consisting of two or more elements in a fixed ratio
- A compound has *emergent properties*, characteristics different from those of its elements

Figure 2.2



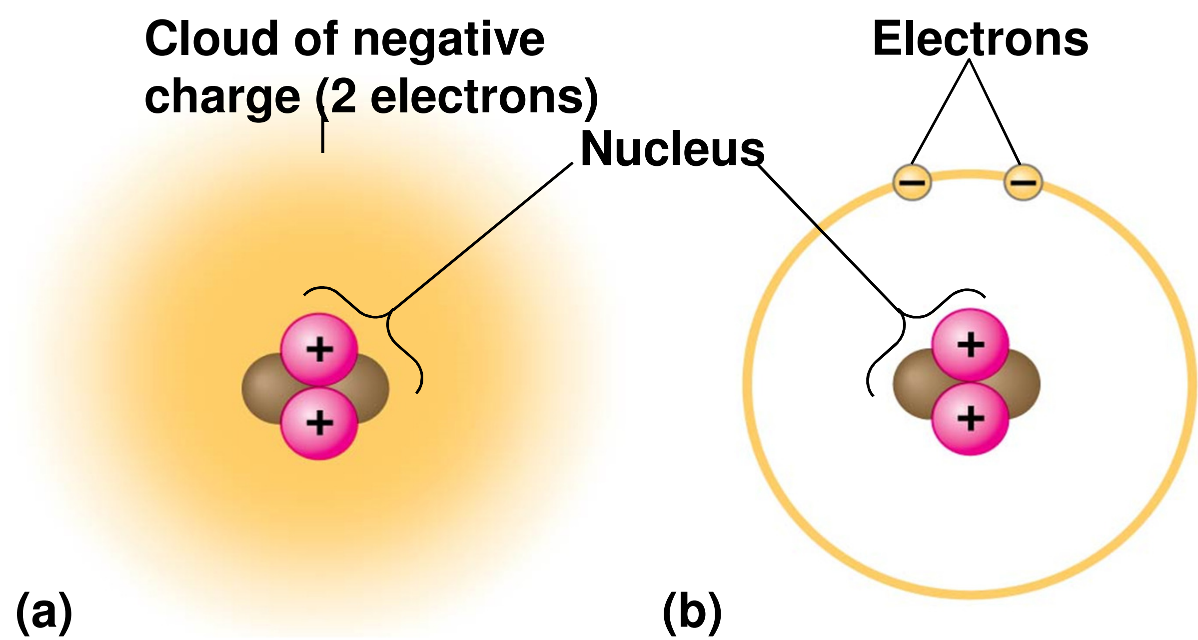
The Elements of Life

- Of 92 natural elements, about 20–25% are **essential elements**, needed by an organism to live a healthy life and reproduce
- **Trace elements** are required in only minute quantities
  - Still considered essential elements!
  - Just needed in smaller quantities
- For example, in vertebrates, iodine (I) is required for normal activity of the thyroid gland
  - In humans, an iodine deficiency can cause goiter

Concept 2.2: An element's properties depend on the structure of its atoms

- Each element consists of a certain type of atom, different from the atoms of any other element
- An **atom** is the smallest unit of matter that still retains the properties of an element
- Atoms are composed of smaller parts called subatomic particles
  - **Neutrons** (no electrical charge)
  - **Protons** (positive charge)
  - **Electrons** (negative charge)

- Neutrons and protons form the **atomic nucleus**
- Electrons form a cloud around the nucleus
- Neutron mass and proton mass are almost identical and are measured in **daltons**



Atomic Number and Atomic Mass

- Atoms of the various elements differ in number of subatomic particles
- An element's **atomic number** is the number of protons in its nucleus
  - In a neutral atom, atomic number also tells us the number of electrons
- An element's **mass number** is the sum of protons plus neutrons in the nucleus
- **Atomic mass**, the atom's total mass, can be approximated by the mass number

Mass number = number of protons + neutrons
= 23 for sodium

$^{23}_{11}\mathrm{Na}$

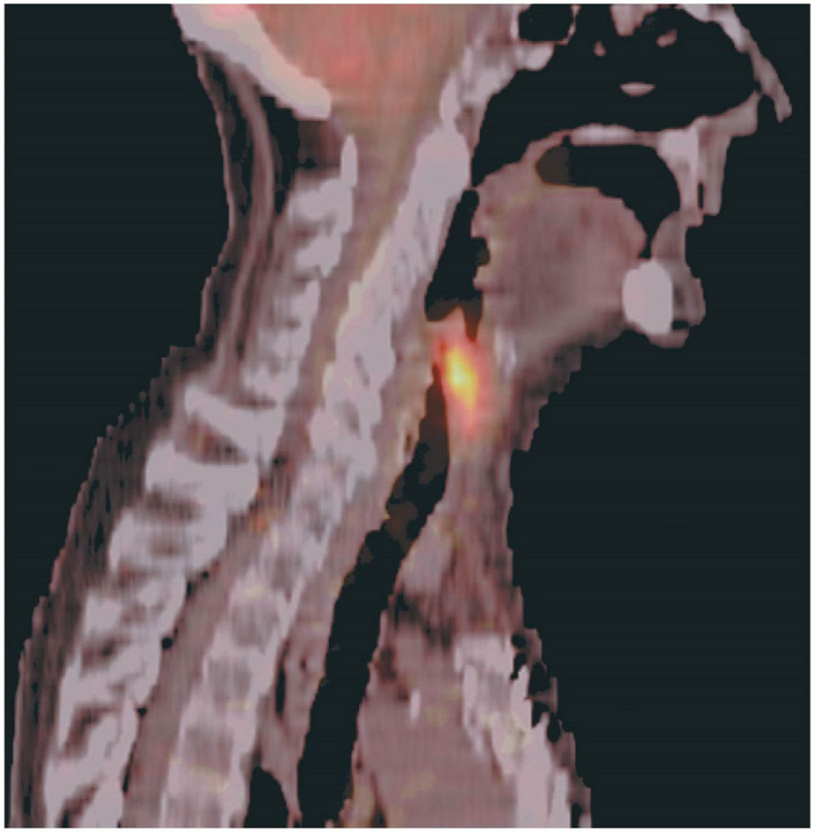
Atomic number = number of protons
= 11 for sodium

Because neutrons and protons each have a mass of approximately 1 dalton, we can estimate the **atomic mass** (total mass of one atom) of sodium as 23 daltons

Isotopes

- All atoms of an element have the same number of protons but may differ in number of neutrons
- **Isotopes** are two atoms of an element that differ in number of neutrons
  - In nature, an element occurs as a mixture of its isotopes
- **Radioactive isotopes** decay spontaneously, giving off particles and energy
  - When the decay leads to a change in number of protons, it transforms the atom to an atom of a different element

- Some applications of radioactive isotopes in biological research are
  - Dating fossils
  - Tracing atoms through metabolic processes
  - Diagnosing medical disorders

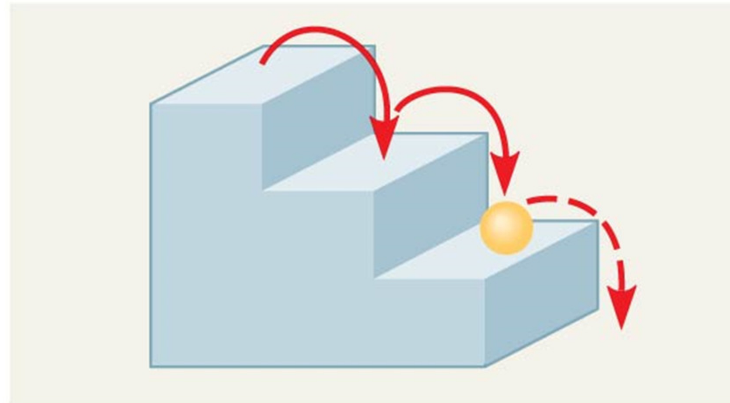


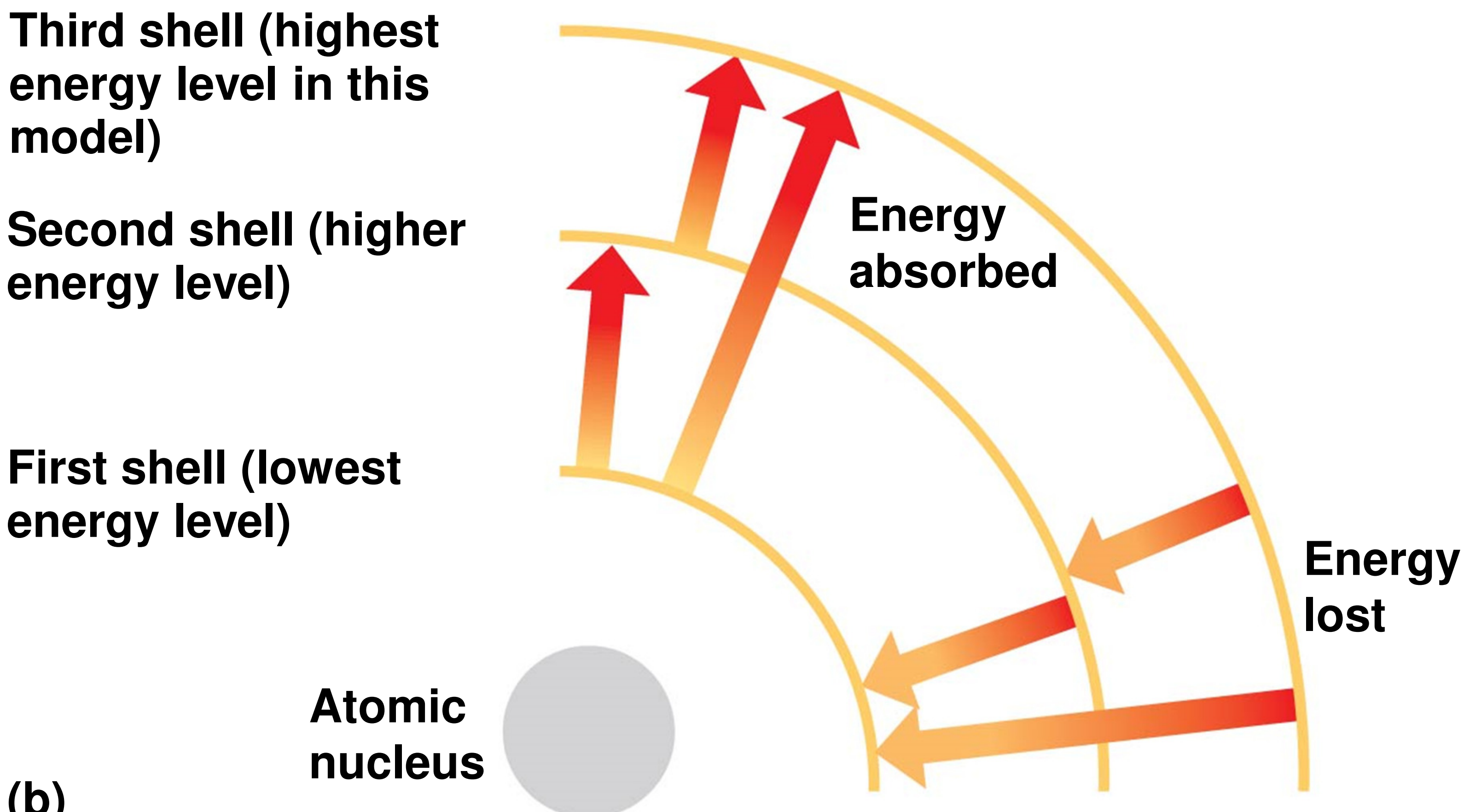
The Energy Levels of Electrons

- **Energy** is the capacity to cause change
  - For instance-to do work
- **Potential energy** is the energy that matter has because of its location or structure
  - Matter has a tendency to move to the lowest possible state of potential energy
- The electrons of an atom differ in their amounts of potential energy based on distance from the nucleus
  - Greater distance from nucleus = greater potential energy
- Changes in potential energy occur in steps of fixed amounts

Figure 2.5

(a) A ball bouncing down a flight of stairs provides an analogy for energy levels of electrons.



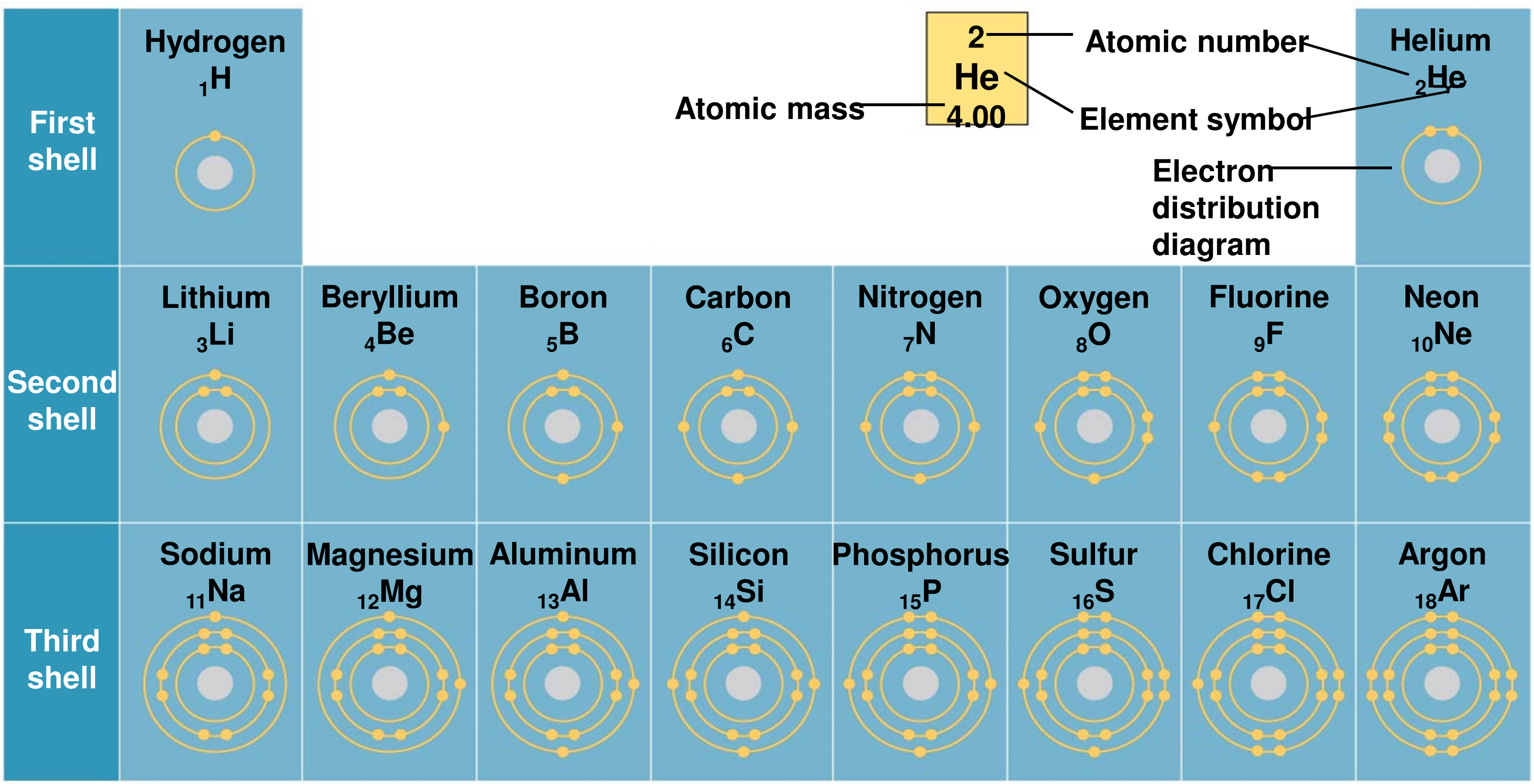


(b)

- Electrons are found in different energy levels, called **electron shells,** each with a characteristic average distance from the nucleus
- The energy level of each shell increases with distance from the nucleus
- Electrons can move to higher or lower shells by absorbing or releasing energy, respectively

Electron Distribution and Chemical Properties

- The chemical behavior of an atom is determined by the distribution of electrons in electron shells
- The periodic table of the elements shows the electron distribution for each element



- Chemical behavior of an atom depends mostly on the number of electrons in its *outermost* shell, or **valence shell**
- **Valence electrons** are those that occupy the valence shell
- Atoms with the same number of valence electrons exhibit similar chemical behavior
- The reactivity of an atom arises from the presence of one or more unpaired electrons in the valence shell
- Atoms with completed valence shells are unreactive, or inert
  - Noble gases = STABLE

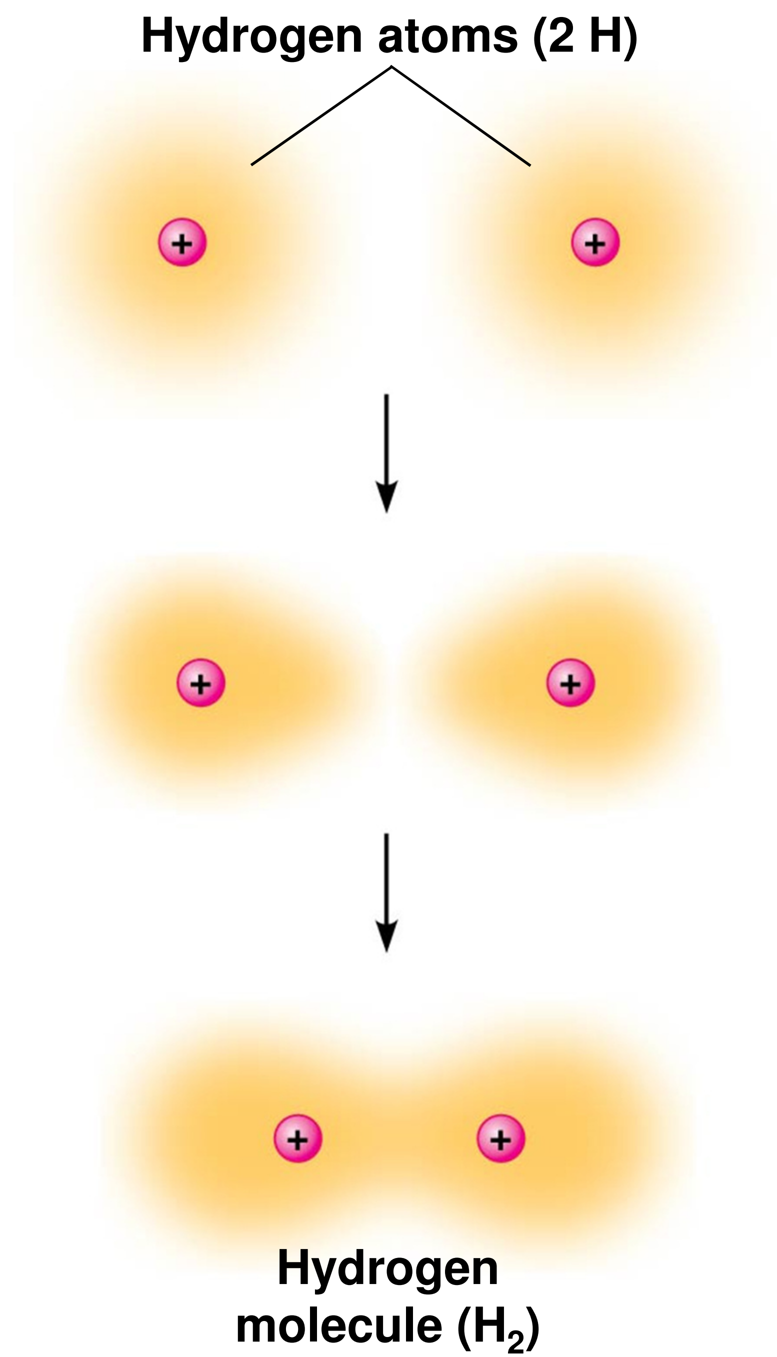
Concept 2.3: The formation and function of molecules depend on chemical bonding between atoms

- Atoms with incomplete valence shells can share or transfer valence electrons with certain other atoms
- This usually results in atoms staying close together, held by attractions called **chemical bonds**

Covalent Bonds

- A **covalent bond** is the *sharing* of a pair of valence electrons by two atoms
- In a covalent bond, the shared electrons count as part of each atom's valence shell
- Two or more atoms held together by valence bonds constitute a **molecule**

Figure 2.7-3



- The notation used to represent atoms and bonding is called a *structural formula*
  - For example, H—H
- This can be abbreviated further with a *molecular formula*
  - For example, H_2

- In a structural formula, a **single bond**, the sharing of one pair of electrons, is indicated by a single line between the atoms
  - For example, H—H

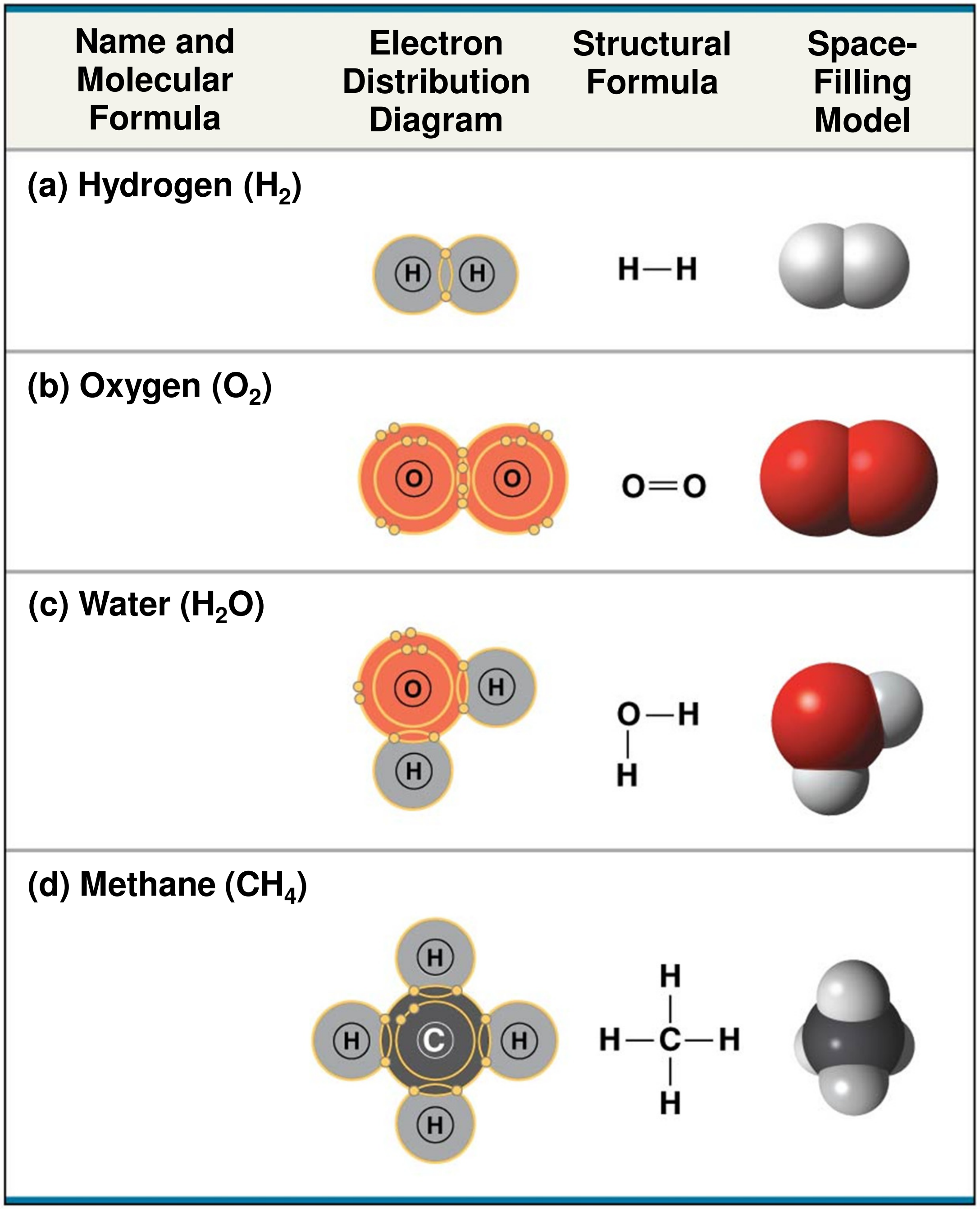
- A **double bond**, the sharing of two pairs of electrons, is indicated by a double line between atoms
  - For example, O=O

- Each atom that can share valence electrons has a bonding capacity, the number of bonds that the atom can form
- Bonding capacity, or **valence**, usually corresponds to the number of electrons required to complete the atom

- Pure *elements* are composed of molecules of <u>one</u> type of atom, such as H_2 and O_2

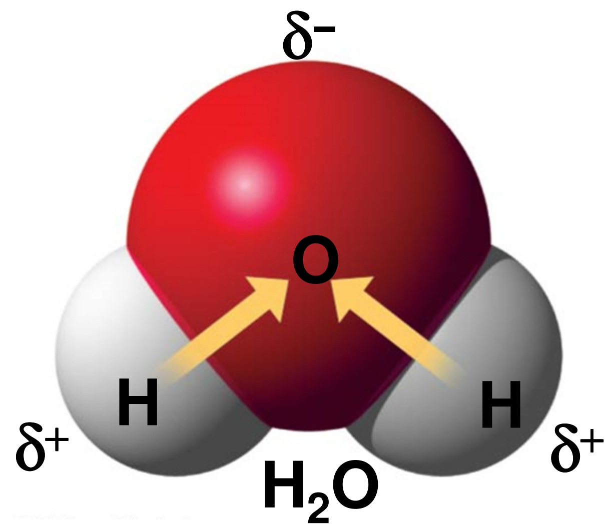
- Molecules composed of a combination of two or more <u>different</u> types of atoms are called *compounds*, such as H_2O or CH_4

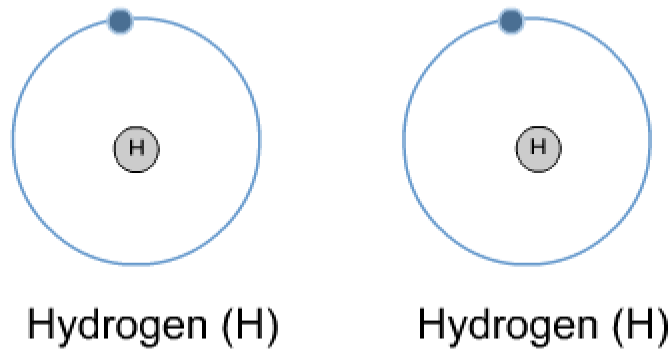
Figure 2.8


Name and Molecular Formula
Electron Distribution Diagram
Structural Formula
Space-Filling Model
(a) Hydrogen (H_2)
H H
H—H
(b) Oxygen (O_2)
O O
O═O
(c) Water (H_2O)
O H H
O—H
|
H
(d) Methane (CH_4)
H H C H H
H
|
H—C—H
|
H

- Atoms in a molecule attract electrons to varying degrees
- **Electronegativity** is an atom's attraction for the electrons in a covalent bond
- The more electronegative an atom, the more strongly it pulls shared electrons toward itself

- In a **nonpolar covalent bond**, the atoms share the electron *equally*
- In a **polar covalent bond**, one atom is more electronegative, and the atoms do *not* share the electron equally
  - Unequal sharing of electrons causes a partial positive or negative charge for each atom or molecule
    - Ex: Water
      - Oxygen atom has a partial negative charge δ-
      - Hydrogen has a partial positive charge δ+



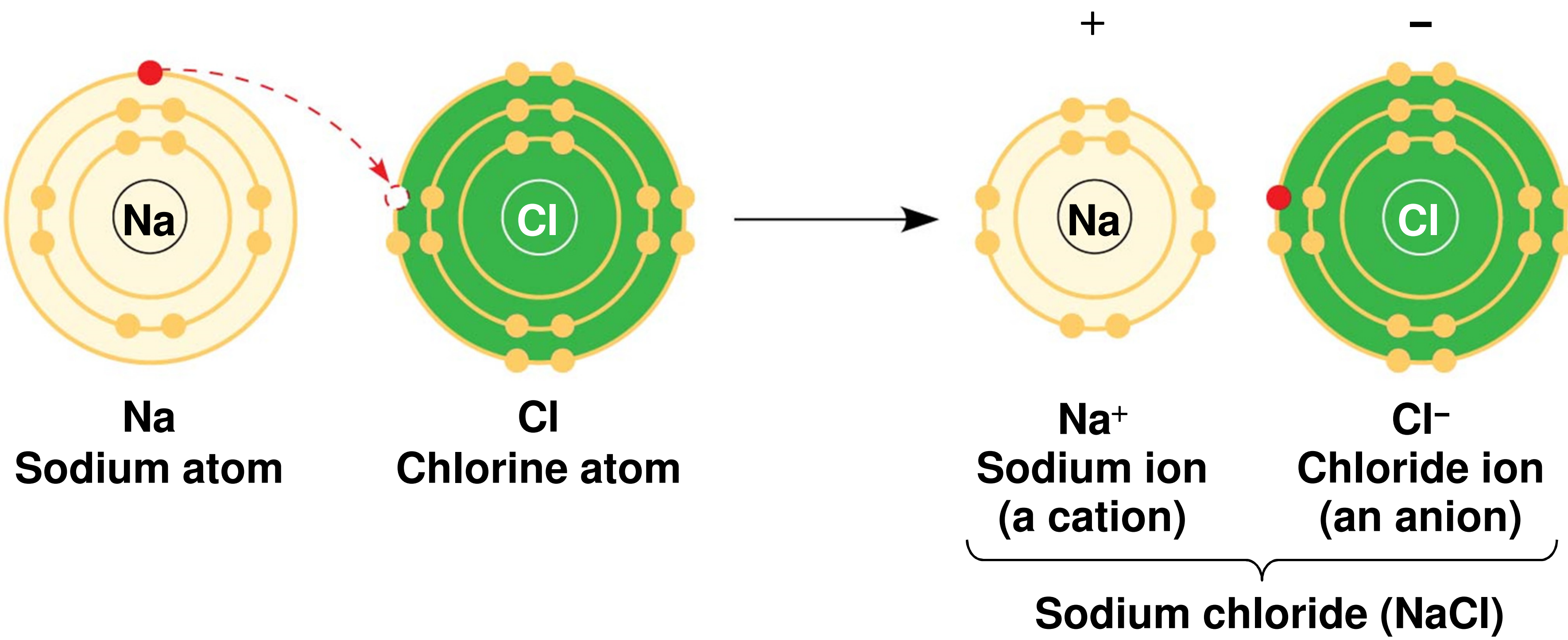


Animation: Covalent Bonds

Ionic Bonds

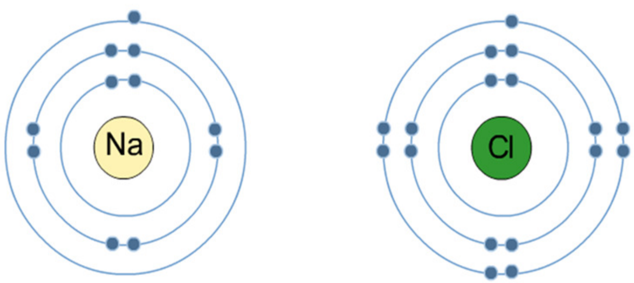
- Atoms sometimes strip electrons from their bonding partners
- An example is the transfer of an electron from sodium to chlorine
- After the transfer of an electron, both atoms have charges
- Both atoms also have complete valence shells

Figure 2.10-2



Na
Cl
+
–
Na
Cl
Na
Sodium atom
Cl
Chlorine atom
Na+
Sodium ion
(a cation)
Cl–
Chloride ion
(an anion)
Sodium chloride (NaCl)

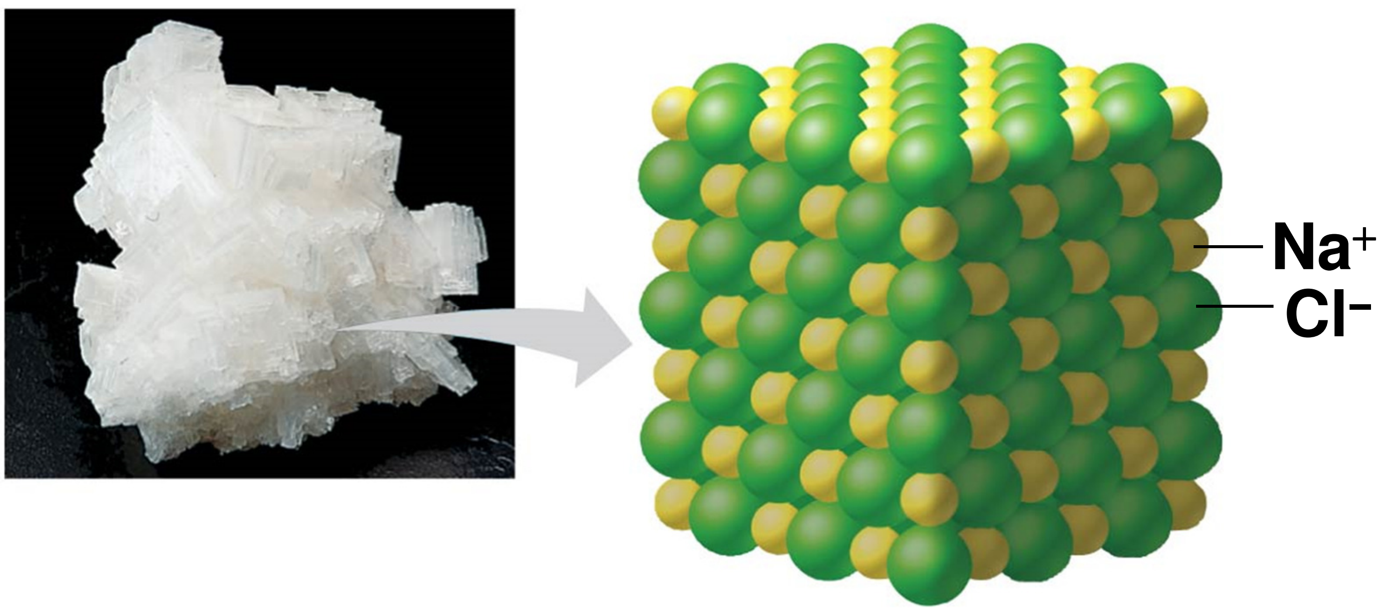
- A charged atom is called an **ion**
- A **cation** is a positively charged ion
  - Less electrons than protons
- An **anion** is a negatively charged ion
  - More electrons than protons
- An **ionic bond** is an attraction between an anion and a cation
  - *Transfer* of electrons

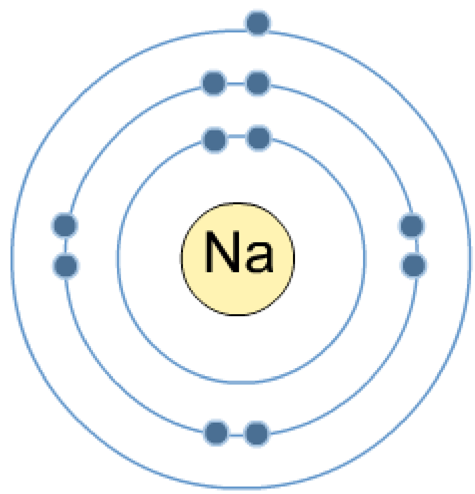


Sodium (Na)
11 protons
11 electrons

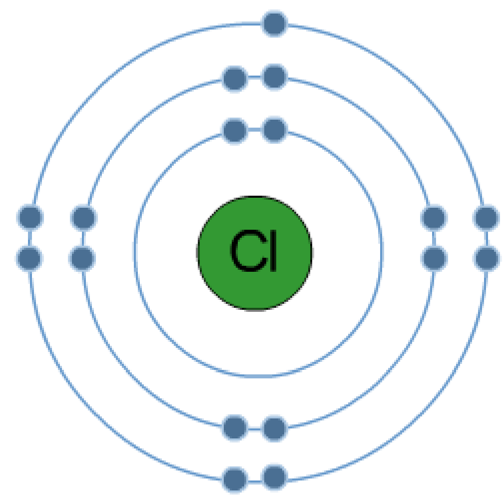
Chlorine (Cl)
17 protons
17 electrons

- Compounds formed by ionic bonds are called **ionic compounds**, or **salts**
- Salts, such as sodium chloride (table salt), are often found in nature as crystals





Sodium (Na)
11 protons
11 electrons



Chlorine (Cl)
17 protons
17 electrons

Animation: Ionic Bonds

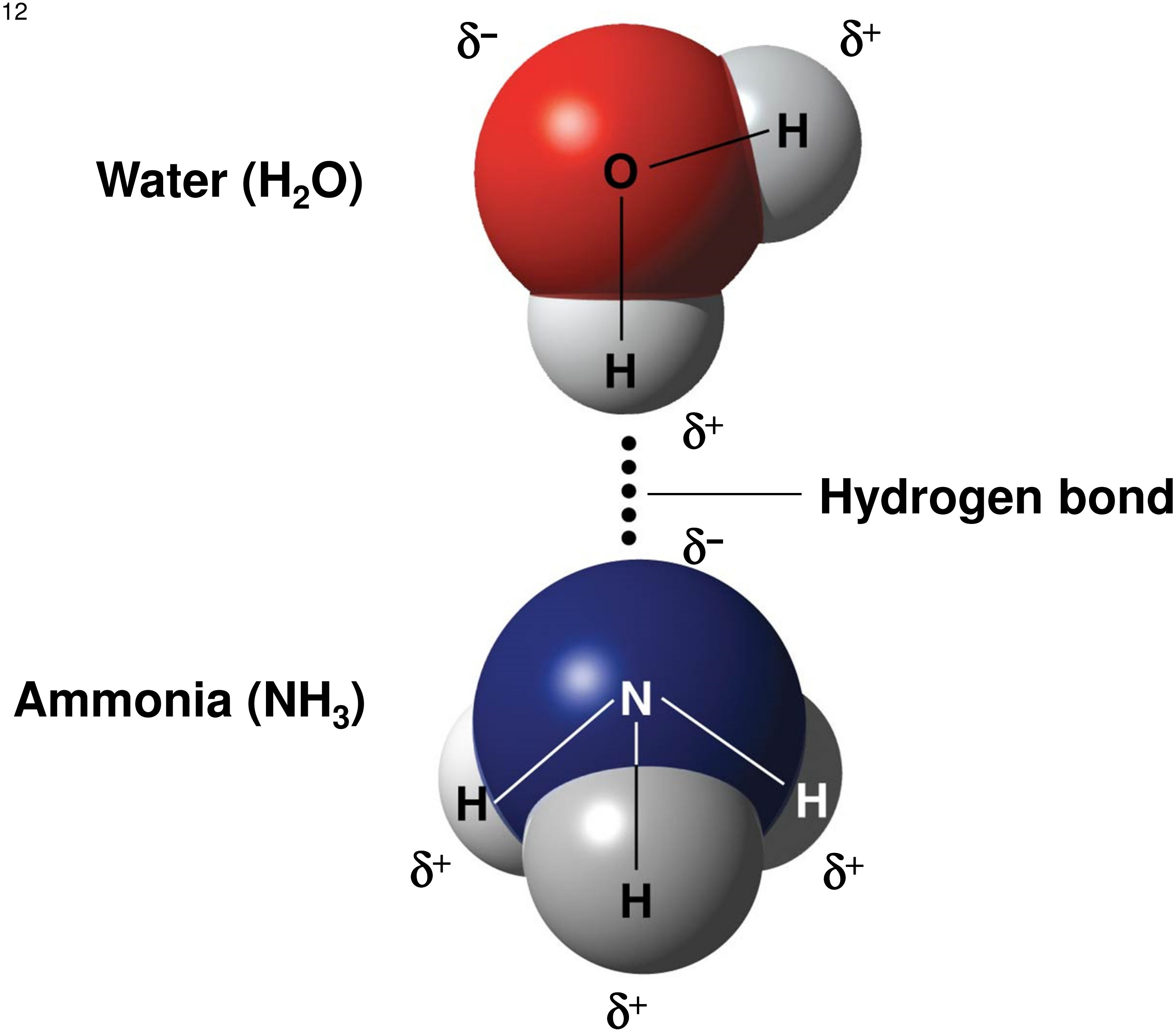
Weak Chemical Bonds

- Most of the strongest bonds in organisms are covalent bonds that form a cell's molecules
- Ionic bonds are also important in organisms
- Weaker bonds, such as hydrogen bonds and van der Waals interactions, are also crucial to life
  - Many large biological molecules are held in their functional form by weak bonds
  - The cumulative effect of weak bonds is to reinforce the 3-D shape of the molecule

Hydrogen Bonds

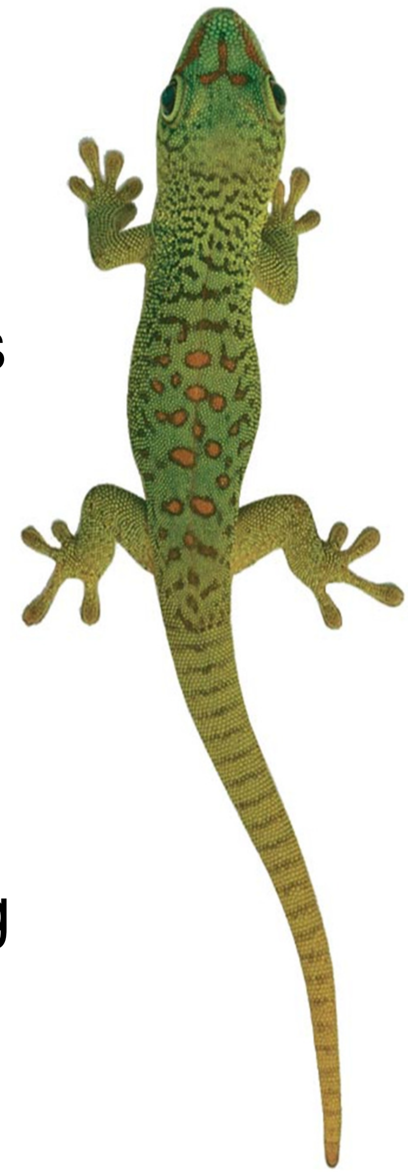
- A **hydrogen bond** forms when a hydrogen atom covalently bonded to one electronegative atom is also attracted to another electronegative atom
- In living cells, the electronegative partners are usually oxygen or nitrogen atoms
- Important biological roles
  - Hydrogen bonds between nitrogenous bases hold one strand of DNA to the other, while still allowing the strands to “unzip” for replication
  - Cohesion of water molecules

Figure 2.12



Van der Waals Interactions

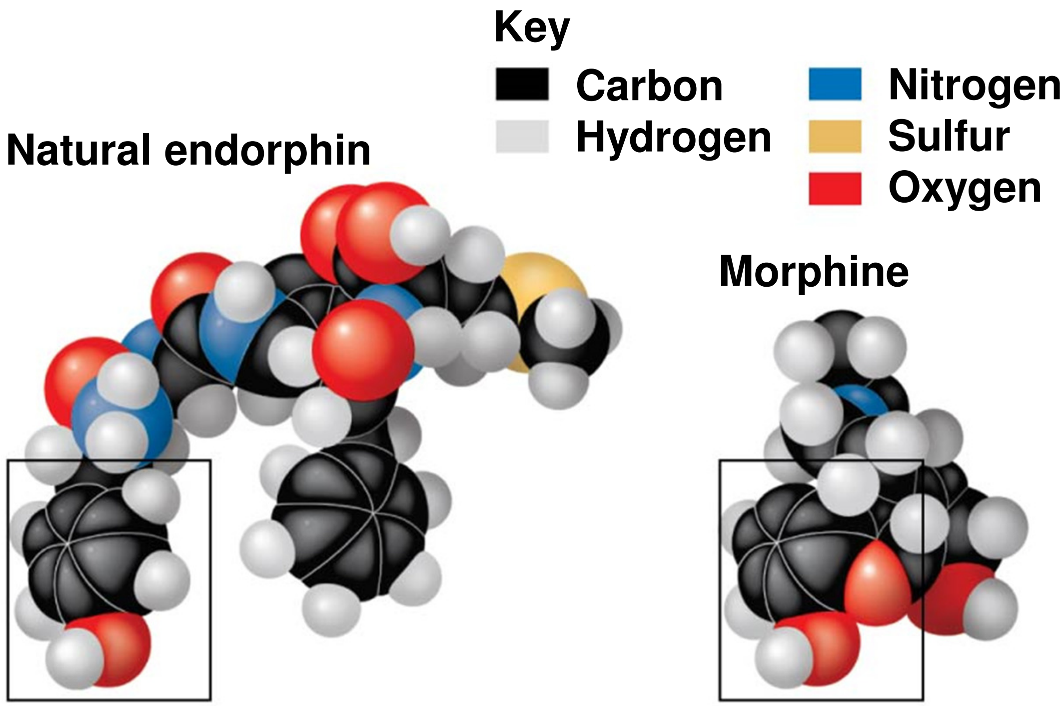
- If electrons are distributed asymmetrically in molecules or atoms, they can result in "hot spots" of positive or negative charge
- **Van der Waals interactions** are attractions between molecules that are close together as a result of these charges
- Van der Waals interactions are individually weak and occur only when atoms and molecules are very close together
- Collectively, such interactions can be strong as between molecules of a gecko's toe hairs and a wall surface



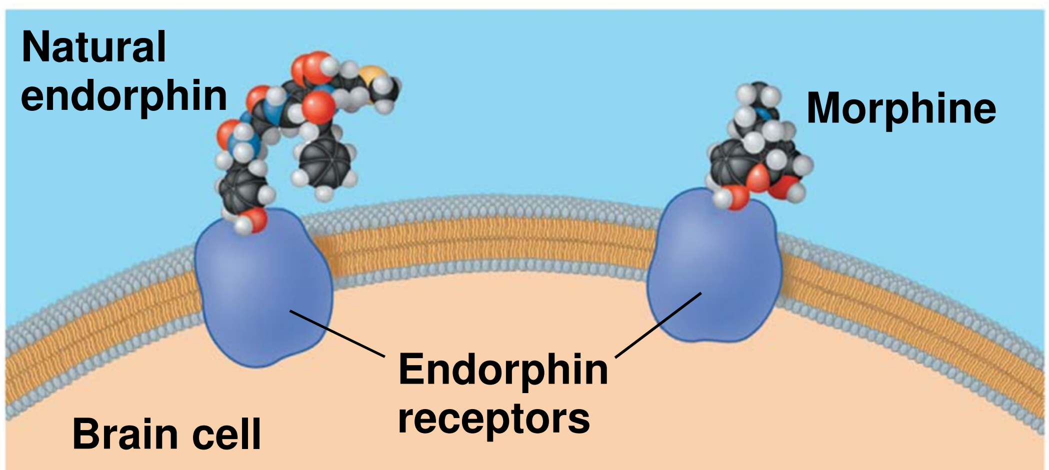
Molecular Shape and Function

- A molecule's *shape* is usually very important to its *function*
- Molecular shape determines how biological molecules recognize and respond to one another
- Biological molecules recognize and interact with each other with a specificity based on molecular shape
- Molecules with similar shapes can have similar biological effects

Figure 2.14



(a) Structures of endorphin and morphine



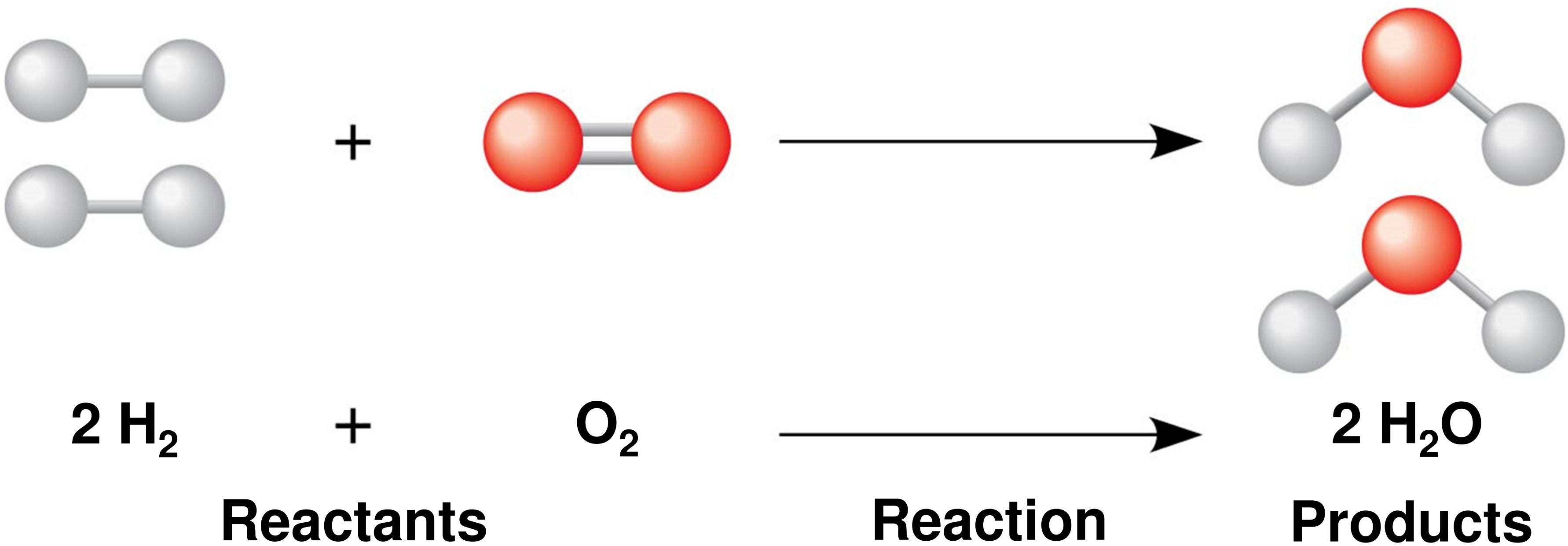
(b) Binding to endorphin receptors

Concept 2.4: Chemical reactions make and break chemical bonds

- **Chemical reactions** are the making and breaking of chemical bonds
- The starting molecules of a chemical reaction are called **reactants**
- The final molecules of a chemical reaction are called **products**

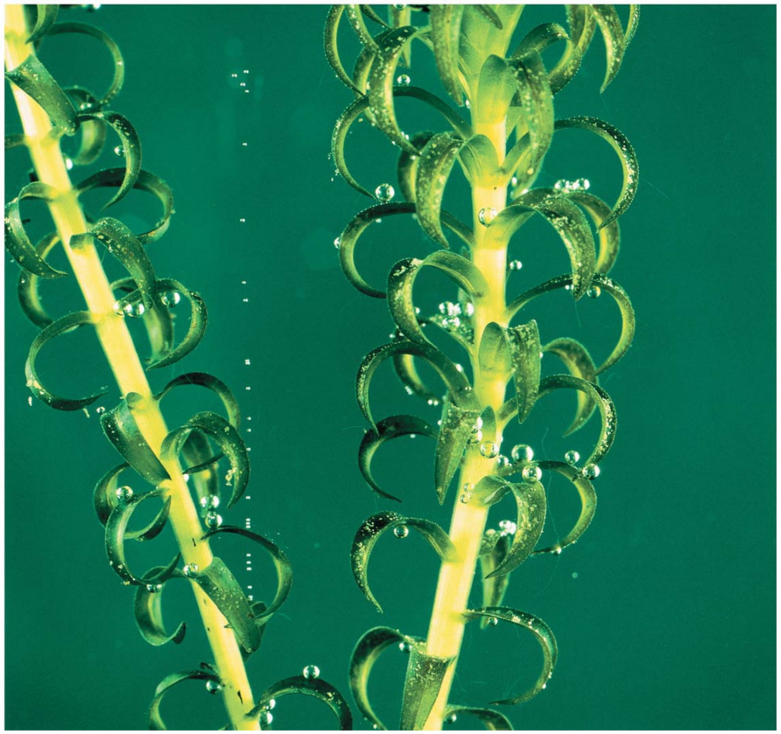
- Matter is conserved in a chemical reaction
  - Reactions cannot create or destroy matter but can only rearrange it!

Figure 2.UN02



- Photosynthesis is an important chemical reaction
- Sunlight powers the conversion of carbon dioxide and water to glucose and oxygen

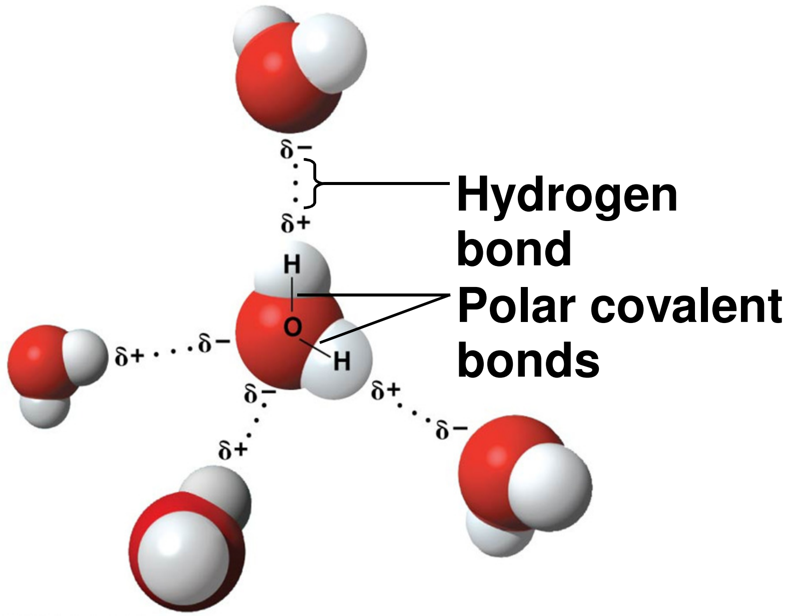
$$6\ CO_2 + 6\ H_2O \rightarrow C_6H_{12}O_6 + 6\ O_2$$



- All chemical reactions are reversible
  - Products of the forward reaction become reactants for the reverse reaction
- One of the factors affecting the rate of a reaction is the concentration of the reactants
- **Chemical equilibrium** is reached when the forward and reverse reaction rates are equal
  - Dynamic equilibrium
    - No net change
    - Concentrations have stabilized at a particular ratio

Concept 2.5: Hydrogen bonding gives water properties that help make life possible on Earth

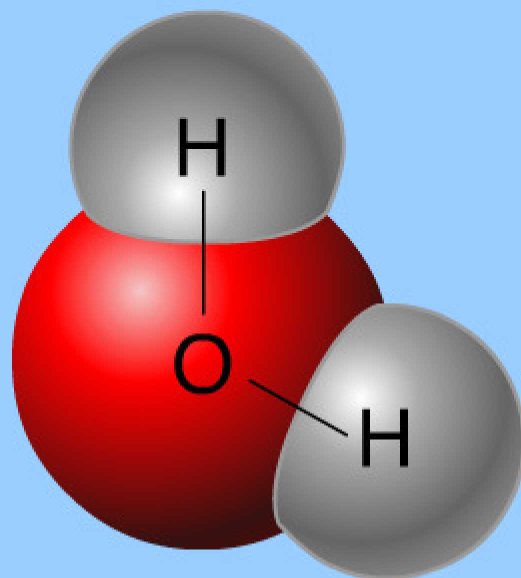
- All organisms are made mostly of water and live in an environment dominated by water
- Water is a **polar molecule**, with the oxygen region having a partial negative charge (δ–) and the hydrogen region a slight positive charge (δ+)
- Properties of water arise from attractions between oppositely charged atoms of different water molecules
- Two water molecules are held together by a hydrogen bond



- Four emergent properties of water contribute to Earth's suitability for life:
  1. Cohesive behavior
  2. Ability to moderate temperature
  3. Expansion upon freezing
  4. Versatility as a solvent

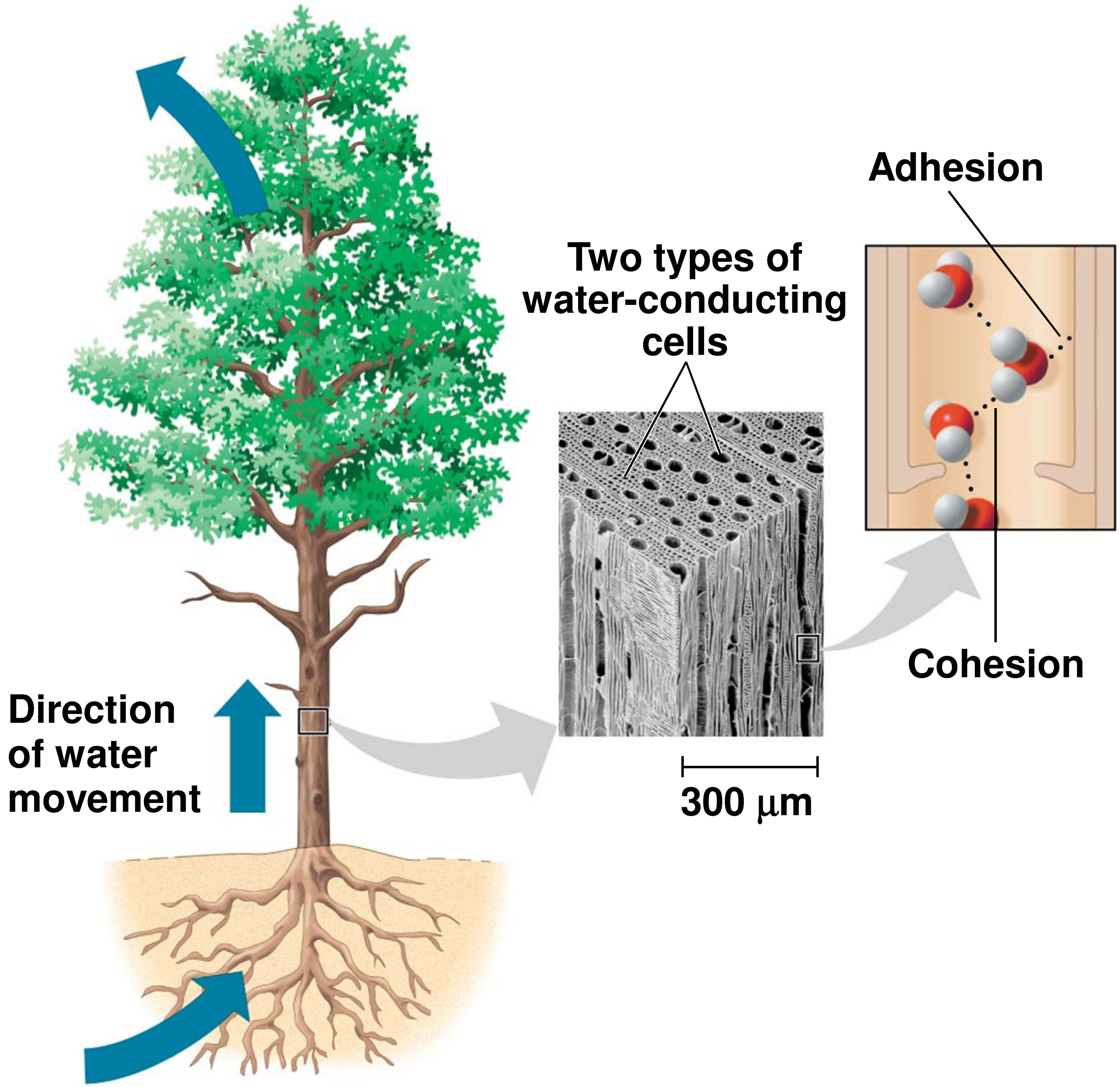
1. Cohesion of Water Molecules

- Water molecules are linked by multiple hydrogen bonds
- The molecules stay close together because of this; it is called **cohesion**
  - Cohesion due to hydrogen bonding contributes to the transport of water and nutrients against gravity in plants
- **Adhesion**, the clinging of one substance to another, also plays a role
  - Adhesion of water to cell walls by hydrogen bonds helps counter the downward pull of gravity

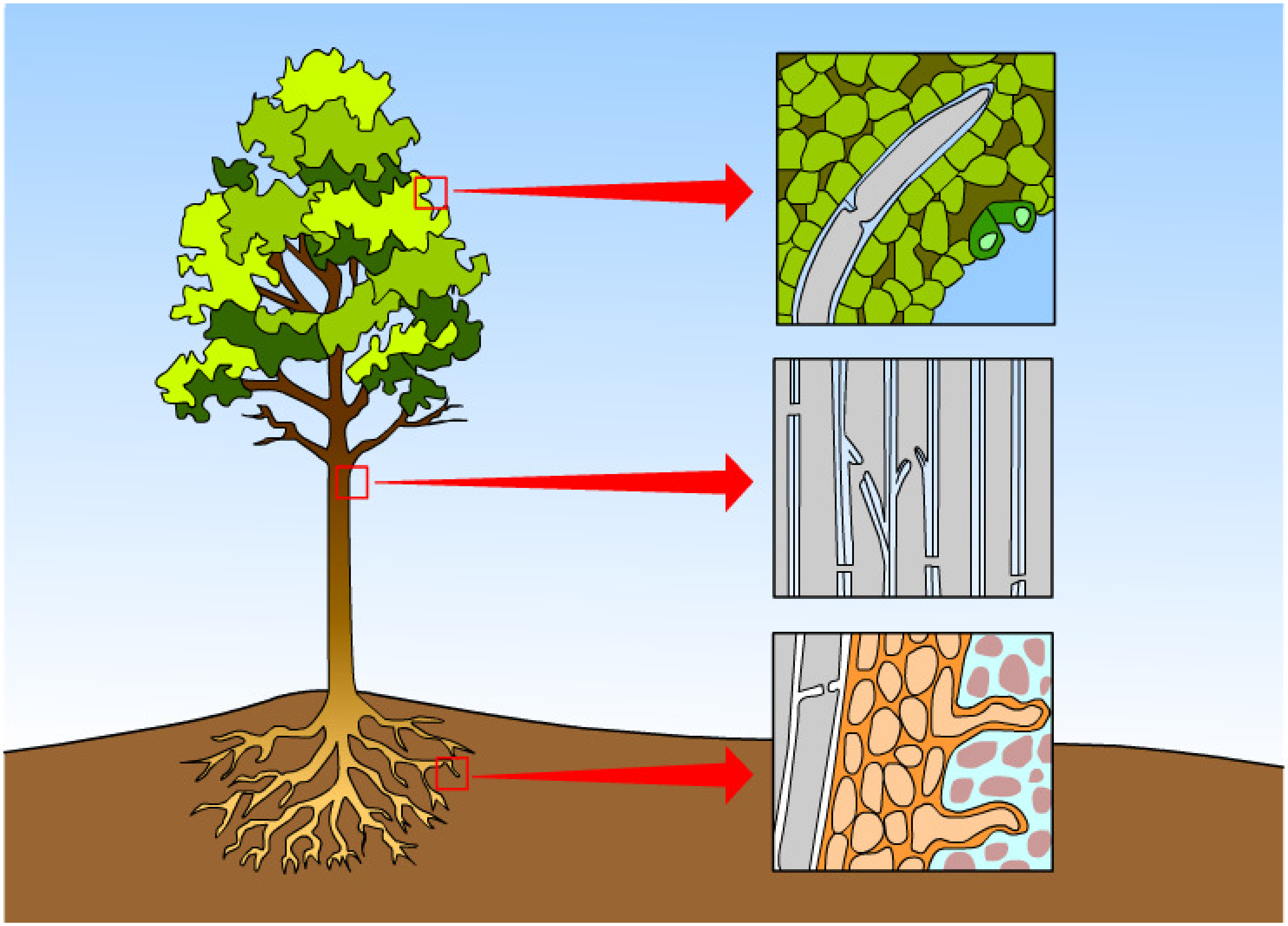


Animation: Water Structure

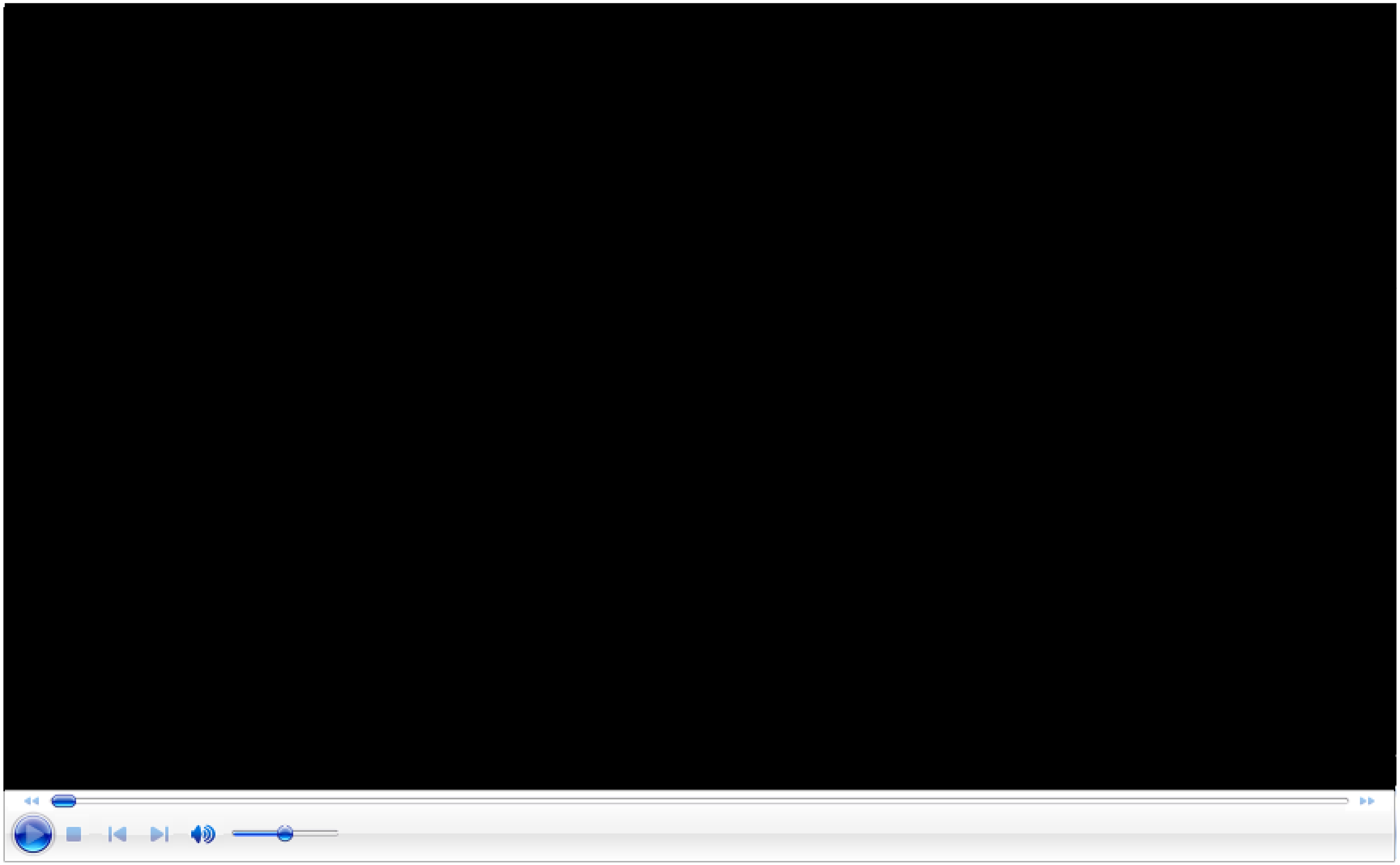
Figure 2.17


Adhesion
Two types of water-conducting cells
Direction of water movement
Cohesion
300 µm

- **Surface tension** is a measure of how hard it is to break the surface of a liquid
- Surface tension is related to cohesion
  - Allows insects and spiders to walk across water without breaking the surface



Animation: Water Transport



Animation: Water Transport in Plants

2. Moderation of Temperature by Water

- Water absorbs heat from warmer air and releases stored heat to cooler air
- Water can absorb or release a large amount of heat with only a slight change in its own temperature

Temperature and Heat

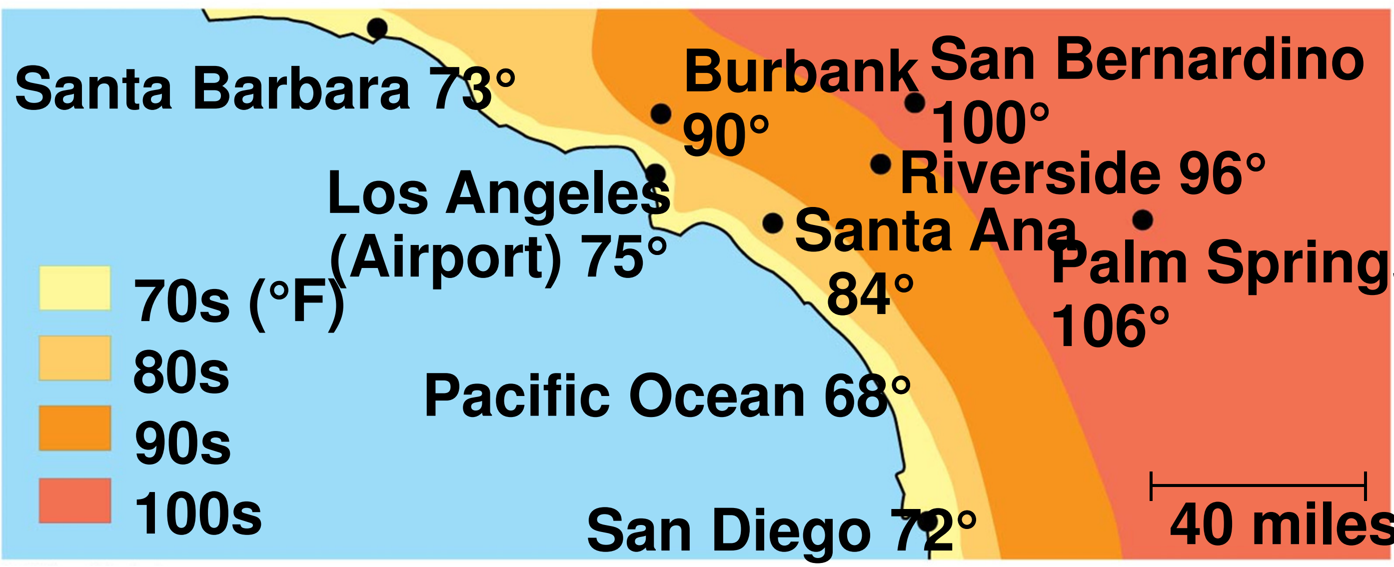
- **Kinetic energy** is the energy of motion
  - The faster a molecules moves, the greater its kinetic energy
- **Thermal energy** is a measure of the total amount of kinetic energy due to molecular motion
- **Temperature** represents the *average* kinetic energy of molecules
- Thermal energy in *transfer* from one body of matter to another is defined as **heat**

- The **Celsius scale** is a measure of temperature using Celsius degrees (°C)
- A **calorie (cal)** is the amount of heat required to raise the temperature of 1 g of water by 1°C
- The "calories" on food packages are actually **kilocalories (kcal)**, where 1 kcal = 1,000 cal
- The **joule (J)** is another unit of energy, where 1 J = 0.239 cal, or 1 cal = 4.184 J

Water's High Specific Heat

- The **specific heat** of a substance is the amount of heat that must be absorbed or lost for 1 g of that substance to change its temperature by 1°C
- The specific heat of water is 1 cal/g/°C
- Water resists changing its temperature because of its high specific heat
  - Because organisms are made primarily of water, they are better able to resist changes in their own temp.

- Water's high specific heat can be traced to hydrogen bonding
  - Heat is absorbed when hydrogen bonds break
  - Heat is released when hydrogen bonds form
- The high specific heat of water keeps temperature fluctuations within limits that permit life

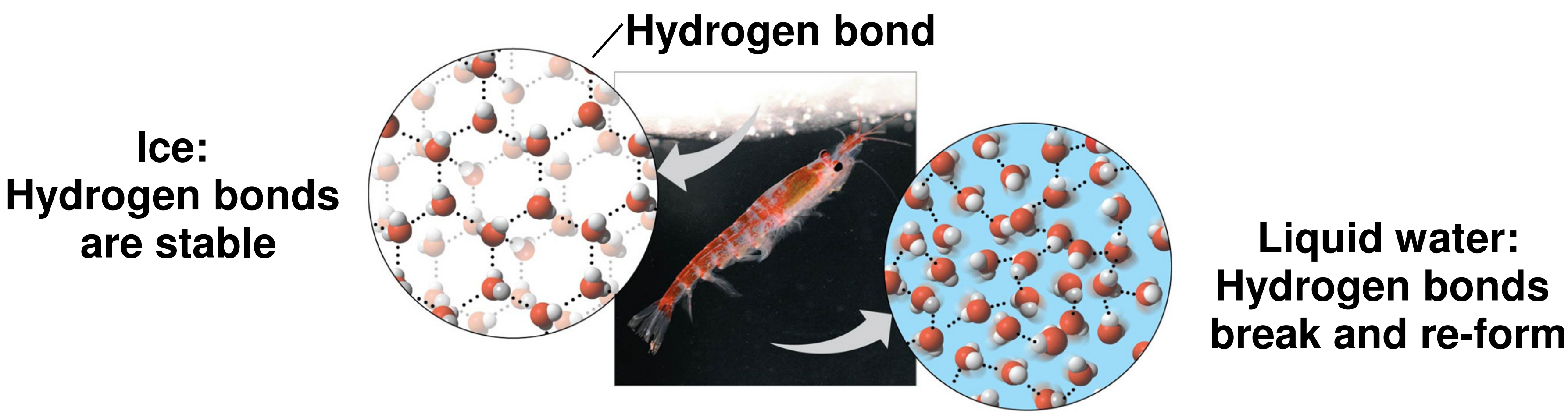


Evaporative Cooling

- Evaporation is transformation of a substance from liquid to gas
- **Heat of vaporization** is the heat a liquid must absorb for 1 g to be converted to gas
  - Water has a high heat of vaporization
  - Helps moderate Earth's climate
- As a liquid evaporates, its remaining surface cools, a process called **evaporative cooling**
  - Evaporative cooling of water helps stabilize temperatures in bodies of water and organisms by preventing overheating

3. Floating of Ice on Liquid Water

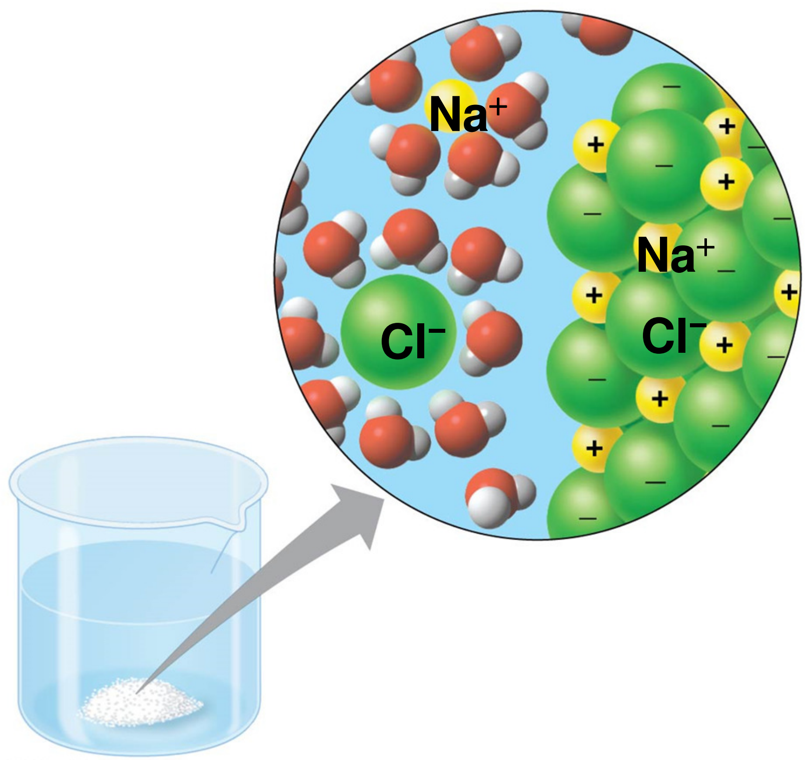
- Ice floats in liquid water because hydrogen bonds in ice are more “ordered,” making ice less dense
  - Water reaches its greatest density at 4°C
- If ice sank, all bodies of water would eventually freeze solid, making life impossible on Earth
  - Ice also insulates liquid water below



4. Water: The Solvent of Life

- A **solution** is a liquid that is a homogeneous mixture of substances
- A **solvent** is the dissolving agent of a solution
- The **solute** is the substance that is dissolved
- An **aqueous solution** is one in which water is the solvent

- Water is a versatile solvent due to its polarity, which allows it to form hydrogen bonds easily
- When an ionic compound is dissolved in water, each ion is surrounded by a sphere of water molecules called a **hydration shell**



Hydrophilic and Hydrophobic Substances

- A **hydrophilic** substance is one that has an affinity for water
  - “Water loving”
- A **hydrophobic** substance is one that does not have an affinity for water
  - “Water fearing”
- Oil molecules are hydrophobic because they have relatively nonpolar bonds

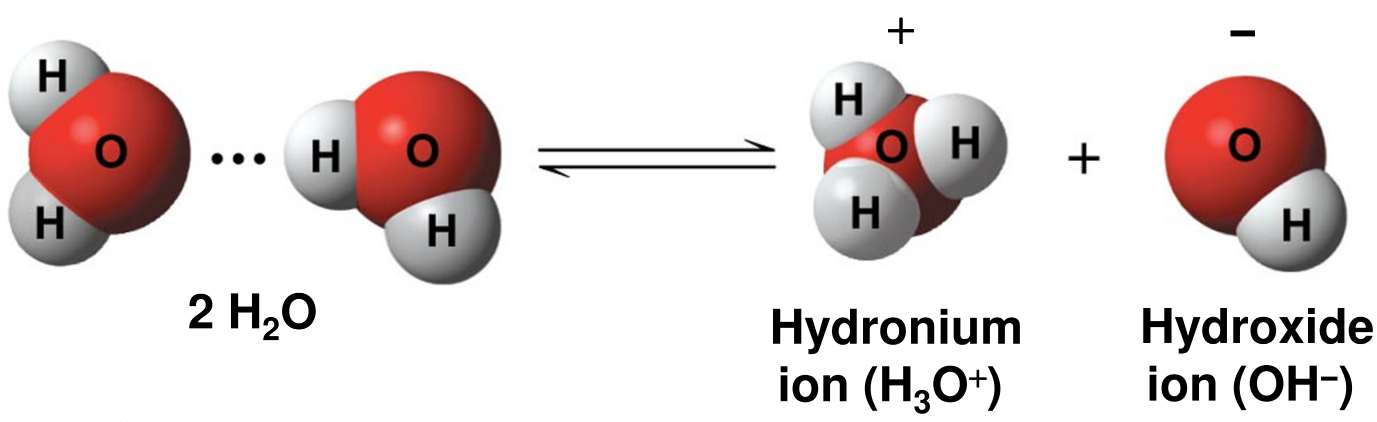
Solute Concentration in Aqueous Solutions

- Most biochemical reactions occur in water
- Chemical reactions depend on collisions of molecules and therefore on the concentration of solutes in an aqueous solution

- **Molecular mass** is the sum of all masses of all atoms in a molecule
- Numbers of molecules are usually measured in moles, where 1 **mole (mol)** = 6.02×10^{23} molecules
- Avogadro's number and the unit *dalton* were defined such that 6.02×10^{23} daltons = 1 g
- **Molarity (*M*)** is the number of moles of solute per liter of solution

Acids and Bases

- Sometimes a **hydrogen ion** (H^+) is transferred from one water molecule to another, leaving behind a **hydroxide ion** (OH^-)
- The proton (H^+) binds to the other water molecule, forming **a hydronium ion** (H_3O^+)
  - By convention, H^+ is used to represent the hydronium ion



- Though water dissociation is rare and reversible, it is important in the chemistry of life
- H^+ and OH^- are very reactive
- Solutes called acids and bases disrupt the balance between H^+ and OH^- in pure water
- **Acids** increase the H^+ concentration in water
- **Bases** reduce the concentration of H^+

- A strong acid like hydrochloric acid, HCl, dissociates completely into H^+ and Cl^- in water:

$$HCl \rightarrow H^+ + Cl^-$$

- Some bases, like sodium hydroxide, NaOH, act as a strong base indirectly by dissociating completely to form hydroxide ions

- These combine with H^+ ions to form water:

$$NaOH \rightarrow Na^+ + OH^-$$

- Some bases, like ammonia, NH_3, act as a relatively weak base when it attracts an H^+ ion from the solution and forms ammonium, NH_4^+
- This is a reversible reaction, as shown by the double arrows:

$$NH_3 + H^+ \rightleftharpoons NH_4^+$$

- Carbonic acid, H_2CO_3, acts as a weak acid, which can reversibly release and accept back H^+ ions:

$$H_2CO_3 \rightleftharpoons HCO_3^- + H^+$$

The pH Scale

- The **pH** of a solution is defined by the negative logarithm of H^+ concentration, written as

$$pH = -\log [H^+]$$

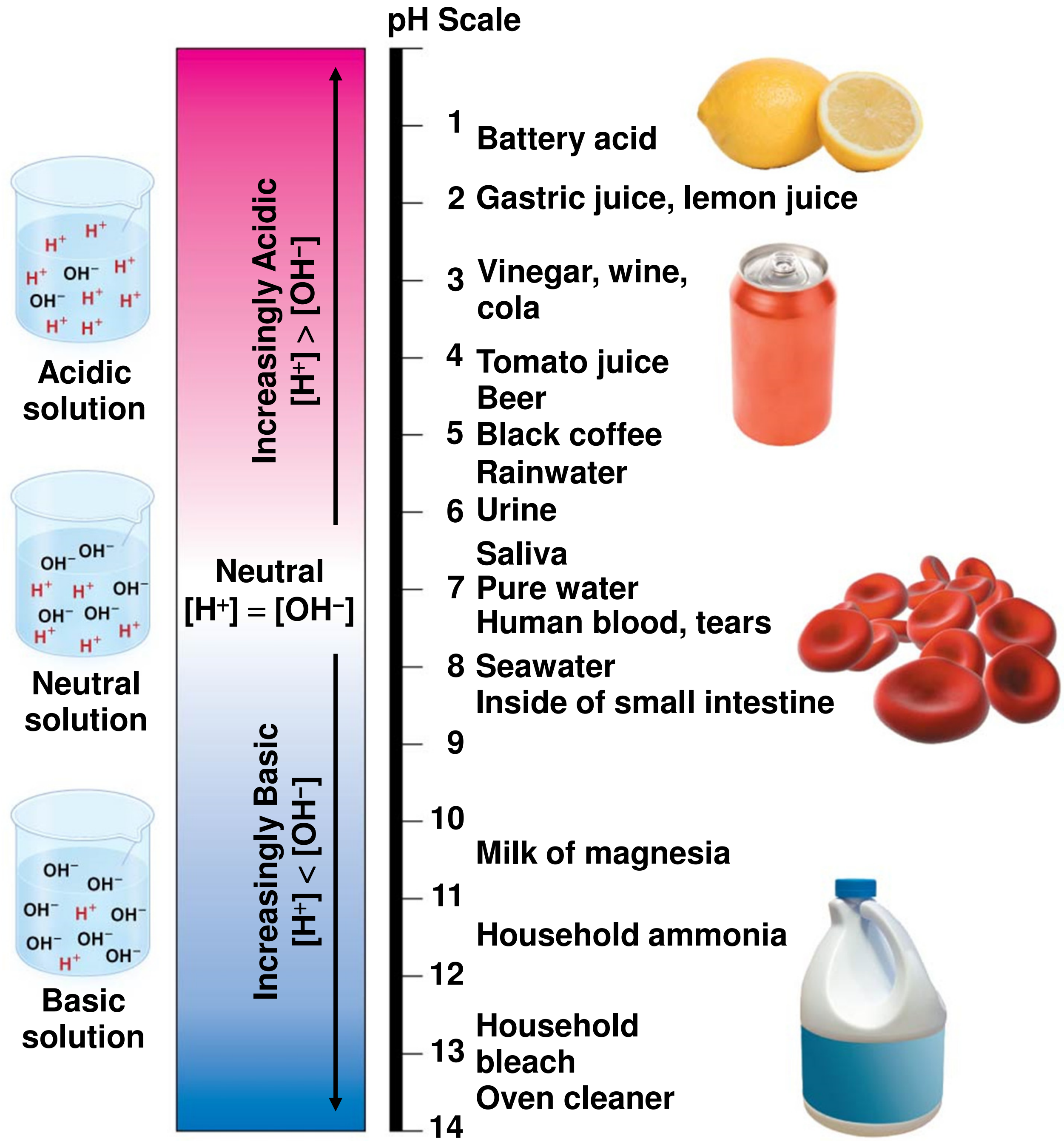
- For a neutral aqueous solution, $[H^+]$ is 10^{-7}, so

$$-\log [H^+] = -\log 10^{-7} = -(-7) = 7$$

- Notice that pH *declines* as H^+ concentration *increases*

- Acidic solutions have pH values less than 7
  - The lower the number, the more acidic the solution
- Basic solutions have pH values greater than 7
- Most biological fluids have pH values in the range of 6 to 8
- Remember, each pH unit represents a TENFOLD difference in H^+ and OH^- concentrations
  - When the pH of a solution changes slightly, the actual concentrations of H^+ and OH^- in the solution change substantially!

Figure 2.23


pH Scale
Increasingly Acidic
[H+] > [OH−]
Neutral
[H+] = [OH−]
Increasingly Basic
[H+] < [OH−]
H+
OH−
Acidic solution
Neutral solution
Basic solution
1 Battery acid
2 Gastric juice, lemon juice
3 Vinegar, wine, cola
4 Tomato juice
Beer
5 Black coffee
Rainwater
6 Urine
Saliva
7 Pure water
Human blood, tears
8 Seawater
Inside of small intestine
9
10
Milk of magnesia
11
Household ammonia
12
Household bleach
13
Oven cleaner
14

Buffers

- The internal pH of most living cells must remain close to pH 7
- **Buffers** are substances that minimize changes in concentrations of H^+ and OH^- in a solution
- Most buffers consist of an acid-base pair that reversibly combines with H^+

- Carbonic acid is a buffer that contributes to pH stability in human blood:

Response to a rise in pH

$$H_2CO_3 \rightleftharpoons HCO_3^- + H^+$$

H_2CO_3	$\rightleftharpoons$	HCO_3^-	+	H^+
H^+ donor (acid)	Response to a drop in pH	H^+ acceptor (base)		Hydrogen ion

Acidification: A Threat to Our Oceans

- Human activities such as burning fossil fuels threaten water quality
- CO_2 is the main product of fossil fuel combustion
- About 25% of human-generated CO_2 is absorbed by the oceans
- CO_2 dissolved in seawater forms carbonic acid; this causes ocean acidification